CHAPTER 3

Matter—Properties and Changes

What You'll Learn

- You will distinguish between physical and chemical properties.
- You will classify matter by composition: element, compound, or mixture.
- You will identify observable characteristics of chemical reactions.
- You will explain the fundamental law of conservation of mass.

Why It's Important

You are completely surrounded by matter. To better understand this matter—how it affects you, how you affect it, and how it can be manipulated for the benefit of society—you need to build a basic understanding of the types and properties of matter.

CLICK HERE

CHEMISTRY Online

Visit the Chemistry Web site at **science.glencoe.com** to find links about matter.

Chemistry is the study of matter and its properties. Every aspect of these divers' environment, under water and on land, is some form of matter.

CONTENTS

DISCOVERY LAB

Materials

large test tube
test-tube holder or rack
10 mL HCl
zinc metal
wood splint
match or burner

Observing Chemical Change

Consider the metal objects that are part of the everyday world. A mailbox, for example, stands outside day in and day out, without seeming to change. Under what conditions does metal exhibit chemical change?

Safety Precautions

Always wear eye goggles, gloves, and an apron when experimenting with chemicals. Use caution when handling an open flame.

Procedure

1. Place a piece of zinc metal in a large test tube.
2. Add approximately 10 mL of 3*M* hydrochloric acid (HCl) to the test tube. Record your observations.
   CAUTION: *HCl causes burns and hazardous fumes.*
3. When the zinc and HCl have reacted for approximately 1 min, bring a lighted, glowing wood splint to the mouth of the test tube. **CAUTION:** *Be sure the test tube is facing away from your face when the splint is brought near.* Again record your observations.

Analysis

What may have caused the dynamic reaction you observed in step 3? Did you expect this reaction? Explain.

Section 3.1

Properties of Matter

Objectives

- **Identify** the characteristics of a substance.
- **Distinguish** between physical and chemical properties.
- **Differentiate** among the physical states of matter.

Vocabulary

substance
physical property
extensive property
intensive property
chemical property
states of matter
solid
liquid
gas
vapor

Imagine yourself scuba diving through a complex biological ecosystem such as a coral reef. What kinds of things fill your imagination? Regardless of what you envision, there is only one answer—you see matter. The diversity of matter in the world and in the universe is astounding. From pepperoni pizzas to supernovas, it's all matter. If we are to understand this diversity, we must start with a way of organizing and describing matter.

Substances

Recall from Chapter 1 that chemistry is the study of matter, and matter is anything that has mass and takes up space. Everything around you is matter; including things such as air and microbes, which you cannot see. For example, table salt is a simple type of matter that you are probably familiar with. Table salt has a unique and unchanging chemical composition. It is always 100% sodium chloride and its composition does not change from one sample to another. Matter that has a uniform and unchanging composition is called a **substance**, also known as a pure substance. Table salt is a substance. Another example of a pure substance is water. Water is always composed of hydrogen and oxygen. Seawater, on the other hand, is not a substance because samples taken from different locations will probably have

Careers Using Chemistry

Science Writer

Do you get excited about news in science and technology? Do you like to explain information in a way that others find interesting and understandable? Then consider a career as a science writer.

Science writers keep up-to-date on what is happening in the world of science and translate that news so nonscientists can understand it. These writers work for newspapers, magazines, scientific publications, television stations, and Internet news services. Lots of curiosity, as well as a degree in a science and/or journalism, is essential.

differing compositions. That is, they will contain differing amounts of water, salts, and other dissolved substances. Given this definition, what other pure substances are you familiar with? Substances are important; much of your chemistry course will be focused on the processes by which substances are changed into different substances.

Physical Properties of Matter

You are used to identifying objects by their properties—their characteristics and behavior. For example, you can easily identify a pencil in your backpack because you recognize its shape, color, weight, or some other property. These characteristics are all physical properties of the pencil. A **physical property** is a characteristic that can be observed or measured without changing the sample's composition. Physical properties describe pure substances, too. Because substances have uniform and unchanging compositions, they have consistent and unchanging physical properties as well. Density, color, odor, taste, hardness, melting point, and boiling point are common physical properties that scientists record as identifying characteristics of a substance. Sodium chloride forms solid, white crystals at room temperature, all having the same unique salty taste. **Table 3-1** lists several common substances and their physical properties.

Table 3-1

Physical Properties of Common Substances

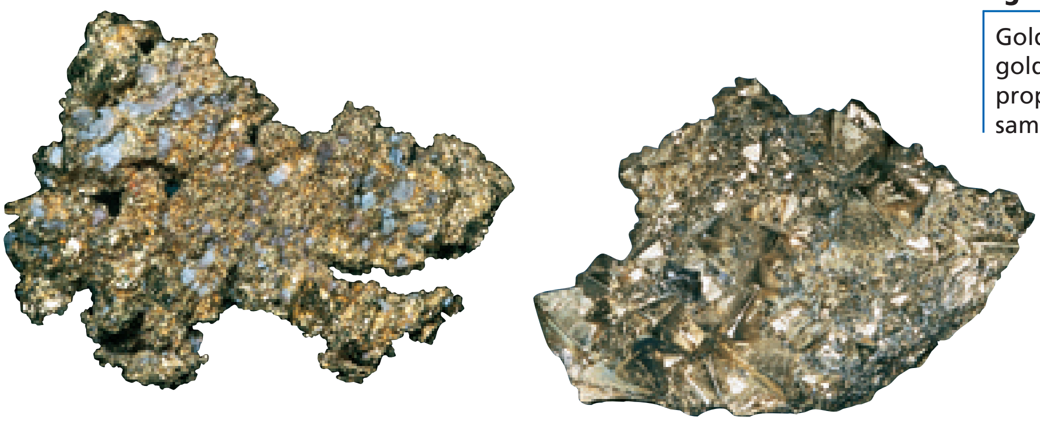
Substance	Color	State at 25°C	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)
Oxygen	Colorless	Gas	−218	−183	0.0014
Mercury	Silver	Liquid	−39	357	13.5
Water	Colorless	Liquid	0	100	1.00
Sucrose	White	Solid	185	Decomposes	1.59
Sodium chloride	White	Solid	801	1413	2.17

Figure 3-1

Miners relied on the physical property of density to distinguish gold (19 g/cm^3) from the worthless minerals in their sluice pans. The density of pyrite, a worthless mineral often mistaken for gold, is 5 g/cm^3.

Extensive and intensive properties Physical properties can be further described as being one of two types. **Extensive properties** are dependent upon the amount of substance present. For example, mass, which depends on the amount of substance there is, is an extensive property. Length and volume are also extensive properties. Density, on the other hand, is an example of an intensive property of matter. **Intensive properties** are independent of the amount of substance present. For example, density of a substance (at constant temperature and pressure) is the same no matter how much substance is present.

A substance can often be identified by its intensive properties. In some cases, a single intensive property is unique enough for identification. During the California gold rush, miners relied on gold's characteristic density (19 g/cm^3) to separate valuable gold-containing flakes from riverbed sand. The process used by the miners is shown in **Figure 3-1**. Another intensive property of gold is its distinctive appearance. Unfortunately, miners often learned that identification of gold based on appearance alone was misleading. **Figure 3-2** shows a nugget of the relatively worthless



a Gold   b Pyrite

Figure 3-2

Gold (a) and pyrite, or "fool's gold" (b), have similar physical properties but are different samples of matter.

mineral pyrite, often called "fool's gold," which looks very similar to actual gold nuggets. Such errors in identification based on the intensive property of appearance fooled many miners into falsely thinking they had struck it rich!

Chemical Properties of Matter

Some properties of a substance are not obvious unless the substance has changed composition as a result of its contact with other substances or the application of thermal or electrical energy. The ability of a substance to combine with or change into one or more other substances is called a **chemical property**. The ability of iron to form rust when combined with air is an example of a chemical property of iron. Similarly, the inability of a substance to change into another substance is also a chemical property. For example, when iron is placed in nitrogen gas at room temperature, no chemical change occurs. The fact that iron does not undergo a change in the presence of nitrogen is another chemical property of iron.

Observing Properties of Matter

Every substance has its own unique set of physical and chemical properties. **Table 3-2** lists several of these properties of copper. **Figure 3-3** shows physical and chemical properties of copper. What physical and chemical properties are evident in these photos?

Table 3-2

Properties of Copper	
Physical properties	**Chemical properties**
• Reddish brown, shiny • Easily shaped into sheets (malleable) and drawn into wires (ductile) • Good conductor of heat and electricity • Density = 8.92 g/cm^3 • Melting point = 1085°C • Boiling point = 2570°C	• Forms green copper carbonate compound when in contact with moist air • Forms new substances when combined with nitric acid and sulfuric acid • Forms a deep blue solution when in contact with ammonia

Figure 3-3

These photos illustrate some of the physical and chemical properties of copper as it exists in the form of hardware (a) and the Statue of Liberty (b).

Observations of properties may vary depending on the conditions of the immediate environment. It is important to state the specific conditions in which observations are made because both chemical and physical properties depend on temperature and pressure. Consider the properties of water, for example. You may think of water as a liquid (physical property) that is not particularly chemically reactive (chemical property). You may also know that water has a density of 1.00 g/cm^3 (physical property). These properties, however, apply only to water at standard "room" temperature and pressure. At temperatures greater than 100°C, water is a gas (physical property) with a density of about 0.0006 g/cm^3 (physical property) that reacts rapidly with many different substances (chemical property). As you can see, the properties of water are dramatically different under different conditions.

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

States of Matter

Imagine you are sitting on a bench, breathing heavily and drinking water after a tiring game of soccer. In this scenario, you are in contact with three different forms of matter; the bench is a solid, the water is a liquid, and the air you breathe is a gas. In fact, all matter that exists on Earth can be classified as one of these physical forms called **states of matter**. Scientists recognize a fourth state of matter called plasma, but it does not occur naturally on Earth except in the form of lightning bolts. The physical state of a substance is a physical property of that substance. Each of the three common states of matter can be distinguished by the way it fills a container.

Solids A **solid** is a form of matter that has its own definite shape and volume. Wood, iron, paper, and sugar are examples of solids. The particles of matter in a solid are very tightly packed; when heated, a solid expands, but only slightly. Because its shape is definite, a solid may not conform to the shape of the container in which it is placed. The tight packing of particles in a solid makes it incompressible; that is, it cannot be pressed into a smaller volume. It is important to understand that a solid is not defined by its rigidity or hardness; the marble statue in **Figure 3-4** is rigid whereas wax sculpture is soft, yet both are solids.

Liquids A **liquid** is a form of matter that flows, has constant volume, and takes the shape of its container. Common examples of liquids include water, blood, and mercury. The particles in a liquid are not rigidly held in place and are less closely packed than are the particles in a solid: liquid particles

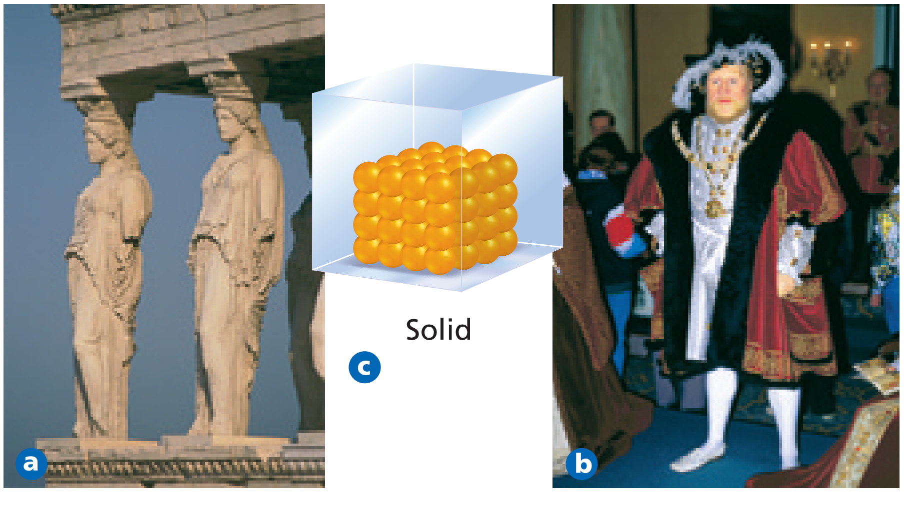


Figure 3-4

The properties of the solid materials marble (a) and wax (b) make these sculptures possible. Particles in a solid are tightly packed (c), giving definite shape and volume to the solid.

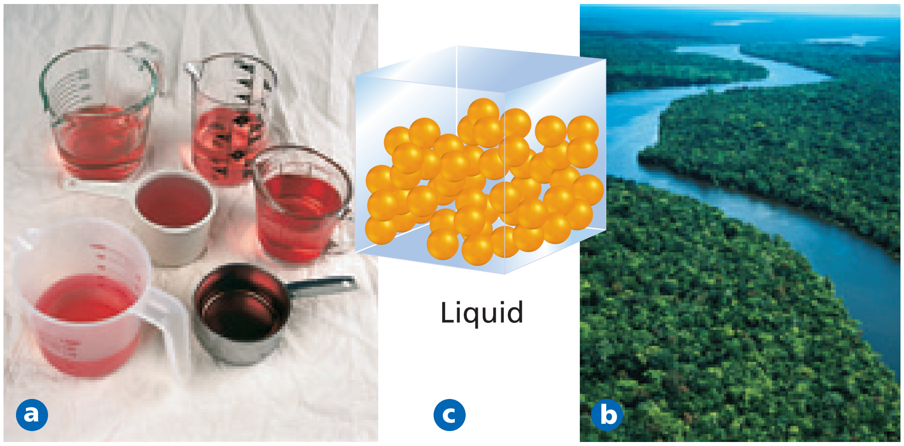


Figure 3-5

a Despite having different shapes, each of these measuring cups holds the same volume of liquid.
b River water flows to fit within the boundaries of its banks, regardless of the curves along its path.
c Molecules in a liquid are closely packed but can still move relatively freely.

are able to move past each other. This allows a liquid to flow and take the shape of its container, although it may not completely fill the container. A liquid's volume is constant: regardless of the size and shape of the container in which the liquid is held, the volume of the liquid remains the same. This is why measuring cups used in cooking, such as those pictured in **Figure 3-5**, can be made in a variety of shapes yet still measure the same volume. Because of the way the particles of a liquid are packed, liquids are virtually incompressible. Like solids, liquids tend to expand when heated.

Gases A **gas** is a form of matter that flows to conform to the shape of its container and fills the entire volume of its container. Examples of gases include neon, which is used in the lighted artwork in **Figure 3-6**; methane, which is used in cooking; and air, which is a mixture of gases. Compared to solids and liquids, the particles of gases are very far apart. Because of the significant amount of space between particles, gases are easily compressed. The **problem-solving LAB** in this section poses several important questions about the practical use of compressed gas.

It is likely that you are familiar with the word vapor as it relates to the word gas. The words gas and vapor, while similar, do not mean the same thing and should not be used interchangeably. The word gas refers to a substance that is naturally in the gaseous state at room temperature. The word **vapor** refers to the gaseous state of a substance that is a solid or a liquid at room temperature. For example, steam is a vapor because at room temperature water exists as a liquid.

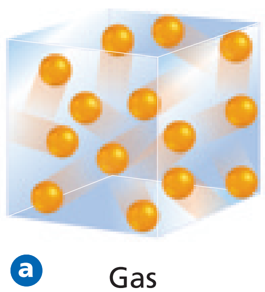


Figure 3-6

a Molecules in a gas are far apart and freely moving.
b Neon gas completely fills the tubes of the electric artwork.

problem-solving LAB

How is compressed gas released?

Recognizing Cause and Effect Tanks of compressed gases are a common sight in a chemistry laboratory. For example, nitrogen is often flowed over a reaction in progress to displace other gases that might interfere with the experiment. Given what you know about the properties of gases, explain how compressed nitrogen is released.

Analysis

By definition, the particles of gases are far apart and gases tend to fill their container, even if the container is a laboratory room. Tanks of compressed gas come from the supplier capped to prevent the gas from escaping. In the lab a chemist or technician attaches a regulator to the tank and secures the tank to a stable fixture.

Thinking Critically

1. Explain why the flow of compressed gas must be controlled for practical use.
2. Predict what would happen if the valve on a full tank of compressed gas were suddenly opened all the way or if the full tank were punctured.

The fact that substances can change form, as in the example of water changing to steam, is another important concept in chemistry. If you review what you just learned about physical properties of substances, you can see that because the particular form of a substance is a physical property, changing the form introduces or adds another physical property to its list of characteristics. In fact, resources that provide tables of physical and chemical properties of substances, such as the *CRC Handbook of Chemistry and Physics,* generally include the physical properties of substances in all of the states in which they can exist.

Section 3.1 Assessment

1. Describe the characteristics that identify a sample of matter as being a substance.
2. Classify each of the following as a physical or chemical property.
   a. Iron and oxygen form rust.
   b. Iron is more dense than aluminum.
   c. Magnesium burns brightly when ignited.
   d. Oil and water do not mix.
   e. Mercury melts at −39°C.
3. Create a table that describes the three common states of matter in terms of their shape, volume, and compressibility.
4. **Thinking Critically** Using what you know about the compressibility of gases, explain why the oxygen in a SCUBA tank is compressed.
5. **Interpreting Data** Bromine is a reddish-brown liquid that boils at 59°C. Bromine is highly reactive with many metals. For example, it reacts with sodium to form a white solid. Classify each of these properties of bromine as either a physical or a chemical property.

Section 3.2 Changes in Matter

You learned in Section 3.1 that scientists can describe matter in terms of physical and chemical properties. For example, a physical property of copper allows it to be drawn into copper wire, and a chemical property of copper accounts for the fact that when a solution of copper ions is combined with ammonia, the copper solution changes to a deep blue color. The key concept in both of these examples is that the substance copper changed in some way. In this section, you'll explore how matter changes as a result of its physical and chemical properties.

Objectives

- **Define** physical change and list several common physical changes.
- **Define** chemical change and list several indications that a chemical change has taken place.
- **Apply** the law of conservation of mass to chemical reactions.

Vocabulary

physical change
chemical change
law of conservation of mass

Physical Changes

A substance often undergoes changes that result in a dramatically different appearance yet leave the composition of the substance unchanged. An example is the crumpling of a sheet of aluminum foil. While the foil goes from a smooth, flat, mirrorlike sheet to a round, compact ball, the actual composition of the foil is unchanged—it is still aluminum. Changes such as this, which alter a substance without changing its composition, are known as **physical changes**. Cutting a sheet of paper and breaking a crystal are other examples of physical changes in matter. Can you name some other physical changes? Your list might include verbs such as bend, grind, crumple, split, and crush, all of which indicate physical change.

As with other physical properties, the state of matter depends on the temperature and pressure of the surroundings. As temperature and pressure change, most substances undergo a change from one state (or phase) to another. For example, at atmospheric pressure and at temperatures below 0°C, water is in its solid state, which is known as ice. As heat is added to the ice, it melts and becomes liquid water. This change of state is a physical change because even though ice and water have very different appearances, their composition is the same. If the temperature of the water increases to 100°C, the water begins to boil and liquid water is converted to steam. Melting and formation of a gas are both physical changes and phase changes. **Figure 3-7** shows condensation, another common phase change. When you encounter terms such as boil, freeze, condense, vaporize, or melt in your study of chemistry, the meaning generally refers to a phase change in matter.

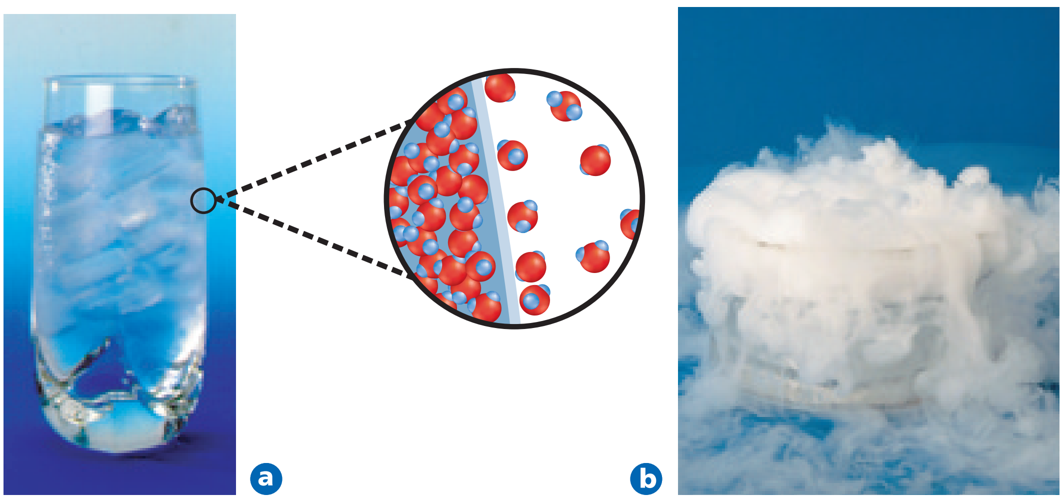


Figure 3-7

a Condensation on an icy beverage glass is the result of the phase change of water in a gaseous state to water in a liquid state.
b The characteristic "fog" of dry ice is actually fine water droplets formed by condensation of water vapor from the air surrounding the very cold dry ice. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

The temperature and pressure at which a substance undergoes a phase change are important physical properties. These properties are listed as the melting and boiling points of the substance. **Table 3-1** on page 56 provides this information for several common substances. Like density, the melting point and boiling point are intensive physical properties that may be used to identify unknown substances. For example, if an unknown solid melts at 801°C and boils at 1413°C—very high temperatures—it is most probably sodium chloride, or common table salt. Tables of intensive properties, such as those given in the *CRC Handbook of Chemistry and Physics,* are indispensable tools in identifying unknown substances from experimental data.

Chemical Changes

As you learned earlier, chemical properties relate to the ability of a substance to combine with or change into one or more substances. A process that involves one or more substances changing into new substances is called a **chemical change**, which is commonly referred to as a chemical reaction. The new substances formed in the reaction have different compositions and different properties from the substances present before the reaction occurred. For example, the crushing of grapes that is part of the wine-making process is a physical change, but the fermentation of the juice, sugars, and other ingredients to wine is a chemical change. The **Chemistry and Society** feature at the end of the chapter describes some interesting consequences of physical and chemical changes in the production of concrete.

Let's consider again the rusting of iron. When a freshly exposed iron surface is left in contact with air, it slowly changes into a new substance, namely, the rust shown in **Figure 3-8a**. The iron reacts with oxygen in the air to form a new substance, rust. Rust is a chemical combination of iron and oxygen. In chemical reactions, the starting substances are called reactants and the new substances that are formed are called products. Thus iron and oxygen are reactants and rust is a product. When you encounter terms such as explode, rust, oxidize, corrode, tarnish, ferment, burn, or rot, the meaning generally refers to a chemical reaction in which reactant substances produce different product substances.

Figure 3-8

a The formation of a gas or solid when reactants mix often indicates that a chemical reaction has taken place. Rust is the result of a chemical reaction. **b** Color changes generally indicate that a chemical reaction has taken place. One example is the color change of tree leaves in the fall.

Figure 3-9

a Energy changes indicate chemical reactions. For example, the burning of wood is a common example of a reaction that releases heat.
b The change in the smell of a substance or the production of an odor may be an indication of a chemical reaction.

Evidence of a chemical reaction As **Figure 3-8a** shows, rust is a brownish-orange powdery substance that looks very different from iron and oxygen. Rust is not attracted to a magnet, whereas iron is. The observation that the product (rust) has different properties than the reactants (iron and oxygen) is evidence that a chemical reaction has taken place. A chemical reaction always produces a change in properties. **Figures 3-8 and 3-9** illustrate several common indicators of chemical change. The **CHEMLAB** at the end of the chapter provides a practical laboratory experience with chemical reactions.

Conservation of Mass

Although chemical reactions have been observed over the course of human history, it was only in the late eighteenth century that scientists began to use quantitative tools to monitor chemical changes. The revolutionary quantitative tool developed at this time was the analytical balance, which was capable of measuring very small changes in mass.

By carefully measuring mass before and after many chemical reactions, it was observed that, although chemical changes occurred, the total mass involved in the reaction remained constant. The constancy of mass in chemical reactions was observed so often that scientists assumed the phenomenon must be true for all reactions. They summarized this observation in a scientific law. The **law of conservation of mass** states that mass is neither created nor destroyed during a chemical reaction—it is conserved. This law was one of the great achievements of eighteenth-century science. The equation form of the law of conservation of mass is

$$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$$

The French scientist Antoine Lavoisier (1743–1794) was one of the first to use an analytical balance like the one shown in **Figure 3-10** to monitor chemical reactions. He studied the thermal decomposition of mercury(II) oxide, known then as calx of mercury. Mercury(II) oxide is a powdery red solid. When it is heated, the red solid reacts to form silvery liquid mercury and colorless oxygen gas as shown in **Figure 3-11** on the next page. The color change and production of a gas are indicators of a

Figure 3-10

The development of scientific tools such as this analytical balance gave a degree of precision to measurements that greatly improved general scientific understanding.

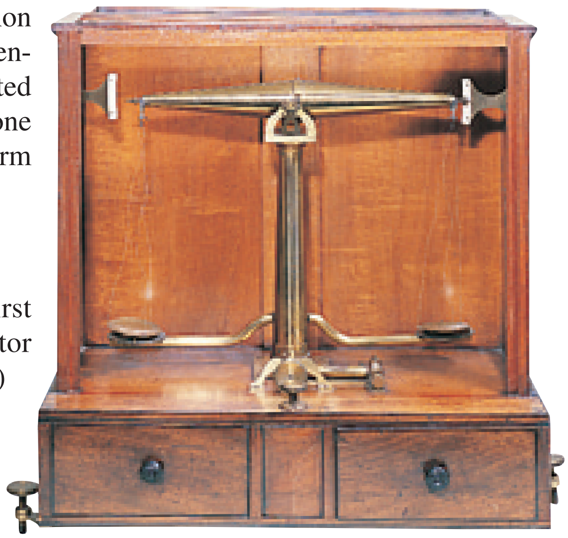
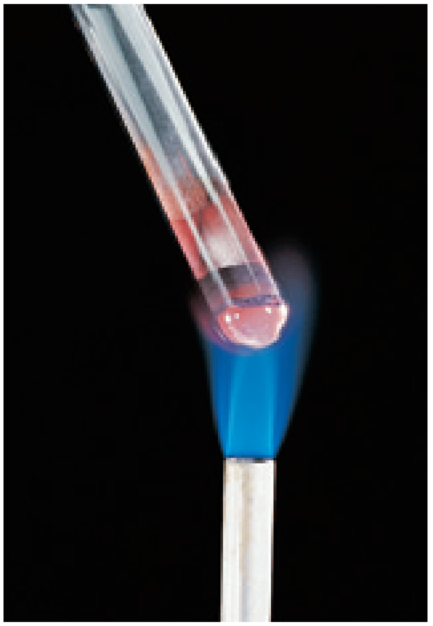


Figure 3-11

Lavoisier's experimental decomposition of mercury(II) oxide is one proof of the law of conservation of mass. Although a chemical reaction is obvious (powder to liquid mercury), matter was neither created nor destroyed.



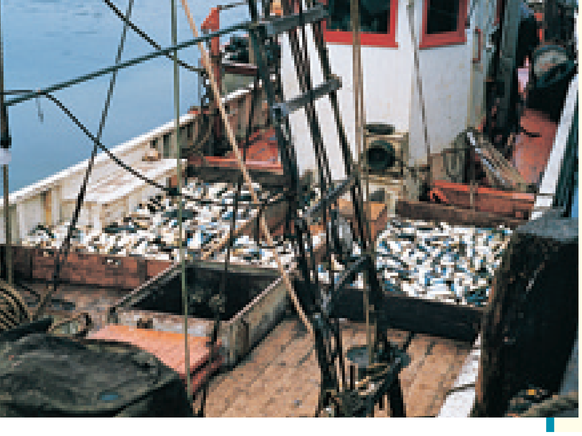
chemical reaction. When the reaction is performed in a closed container, the oxygen gas cannot escape and the mass before and after the reaction can be measured. The masses will be the same.

$$\text{Mercury(II) oxide yields mercury} + \text{oxygen}$$

$$\underset{\text{Mass of reactant}}{216\text{ g}} = \underset{\text{mass of products}}{200\text{ g} + 16\text{ g}}$$

A more modern digital analytical balance can be used to prove the conservation of mass of this example. The law of conservation of mass is one of the most fundamental concepts of chemistry. Let's examine more closely some situations that illustrate the concept. Example Problem 3-1 leads you through a sample calculation. The practice problems also illustrate the law of conservation of mass.

Mercury occurs naturally in air, water, soil, and living organisms. Seafood that is intended for human consumption is monitored to ensure that the products do not contain levels of mercury exceeding the established limits for public safety.



EXAMPLE PROBLEM 3-1

Conservation of Mass

In an experiment, 10.00 g of red mercury(II) oxide powder is placed in an open flask and heated until it is converted to liquid mercury and oxygen gas. The liquid mercury has a mass of 9.26 g. What is the mass of oxygen formed in the reaction?

1. Analyze the Problem

You are given the mass of a reactant and the mass of one of the products in a chemical reaction. Applying the law of conservation of mass, the total mass of the products must equal the total mass of the reactants. This means that the mass of the liquid mercury plus the mass of the oxygen gas must equal the mass of the mercury(II) oxide powder.

Known	Unknown
Mass of mercury(II) oxide = 10.00 g	Mass of oxygen formed = ? g
Mass of liquid mercury = 9.26 g	

2. Solve for the Unknown

Write an equation showing conservation of mass of reactants and products.

$$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$$

$$\text{Mass}_{\text{mercury(II) oxide}} = \text{Mass}_{\text{mercury}} + \text{Mass}_{\text{oxygen}}$$

Solve the equation for $\text{Mass}_{\text{oxygen}}$.

$$\text{Mass}_{\text{oxygen}} = \text{Mass}_{\text{mercury(II) oxide}} - \text{Mass}_{\text{mercury}}$$

Substitute known values and solve.

$$\text{Mass}_{\text{oxygen}} = 10.00 \text{ g} - 9.26 \text{ g}$$

$$\text{Mass}_{\text{oxygen}} = 0.74 \text{ g}$$

3. Evaluate the Answer

The sum of the masses of the two products equals the mass of the reactant, verifying that mass has been conserved. The answer is correctly expressed to the hundredths place.

PRACTICE PROBLEMS

6. From a laboratory process designed to separate water into hydrogen and oxygen gas, a student collected 10.0 g of hydrogen and 79.4 g of oxygen. How much water was originally involved in the process?

7. A student carefully placed 15.6 g of sodium in a reactor supplied with an excess quantity of chlorine gas. When the reaction was complete, the student obtained 39.7 g of sodium chloride. How many grams of chlorine gas reacted? How many grams of sodium reacted?

8. In a flask, 10.3 g of aluminum reacted with 100.0 g of liquid bromine to form aluminum bromide. After the reaction, no aluminum remained, and 8.5 grams of bromine remained unreacted. How many grams of bromine reacted? How many grams of compound were formed?

9. A 10.0-g sample of magnesium reacts with oxygen to form 16.6 g of magnesium oxide. How many grams of oxygen reacted?

For more practice with conservation of mass, go to **Supplemental Practice Problems** in Appendix A.

Section 3.2 Assessment

10. Describe the results of a physical change and list three examples of physical change.

11. Describe the results of a chemical change. List four indicators of chemical change.

12. Solve each of the following.

- **a.** In the complete reaction of 22.99 g of sodium with 35.45 g of chlorine, what mass of sodium chloride is formed?
- **b.** A 12.2-g sample of X reacts with a sample of Y to form 78.9 g of XY. What is the mass of Y that reacted?

13. Thinking Critically A friend tells you, "Because composition does not change during a physical change, the appearance of a substance does not change." Is your friend correct? Explain why.

14. Classifying Classify each of the following examples as a physical change or a chemical change.

- **a.** crushing an aluminum can
- **b.** recycling used aluminum cans to make new aluminum cans
- **c.** aluminum combining with oxygen to form aluminum oxide

Section 3.3 Mixtures of Matter

Objectives

- **Contrast** mixtures and substances.
- **Classify** mixtures as homogeneous or heterogeneous.
- **List** and **describe** several techniques used to separate mixtures.

Vocabulary

mixture
heterogeneous mixture
homogeneous mixture
solution
filtration
distillation
crystallization
chromatography

When scientists speak of the composition of matter, they are referring to the kinds and amounts of components of which the matter is made. On the basis of composition alone, all matter can be classified into two broad categories: substances or mixtures. You have already learned that a pure substance is a form of matter with a uniform and unchanging composition. You also know that the intensive properties of pure substances do not change, regardless of the physical state or amount of the substance. But what is the result when two or more substances are combined?

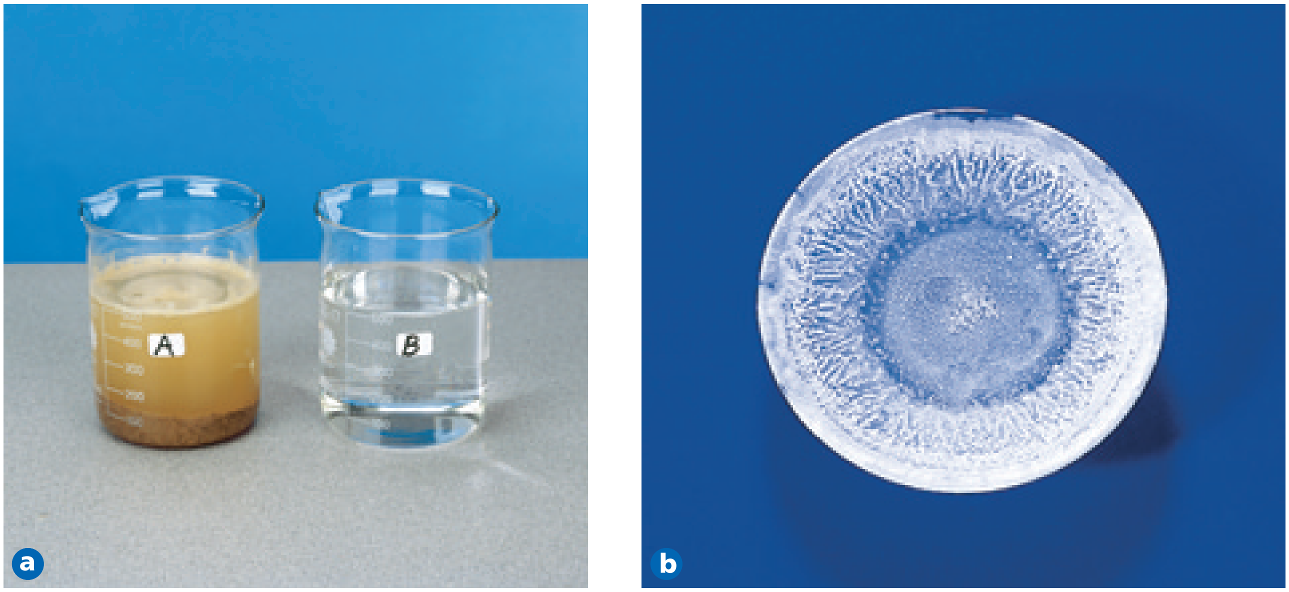
Mixtures

A **mixture** is a combination of two or more pure substances in which each pure substance retains its individual chemical properties. The composition of mixtures is variable, and the number of mixtures that can be created by combining substances is infinite. Although much of the focus of chemistry is the behavior of substances, it is important to remember that most everyday matter occurs as mixtures. Substances tend naturally to mix; it is difficult to keep things pure.

Two mixtures, sand and water, and table salt and water, are shown in **Figure 3-12a**. You know water to be a colorless liquid. Sand is a grainy solid that does not dissolve in water. When sand and water are mixed, the two substances are in contact, yet each substance retains its properties. The sand and water have not reacted. Just by looking at the sand–water mixture in beaker A, it is easy to see each separate substance. Some mixtures, however, may not look like mixtures at all. The mixture of table salt and water in the beaker labeled B is colorless and appears the same as pure water. How can you determine if it is a mixture? If you were to boil away the water, you would see a white residue. That residue, shown in **Figure 13-12b,** is the salt. Thus, the colorless mixture actually contained two separate substances. The salt and the water physically mixed but did not react and were separated by the physical method of boiling.

Figure 3-12

a The components of the sand and water mixture (left) are obvious, whereas the components of the table salt and water mixture (right) are not.
b The salt component becomes obvious when the mixture is boiled.



Types of mixtures The combinations of pure substances shown in **Figure 3-12** are indeed both mixtures, despite their obvious visual differences. Can you think of some way to further define mixtures? Mixtures themselves are classified as either heterogeneous or homogeneous. A **heterogeneous mixture** is one that does not blend smoothly throughout and in which the individual substances remain distinct. The sand and water mixture is an example of a heterogeneous mixture. Suppose you draw a drop from the top of the mixture using an eyedropper. The drop would be almost completely water. If you draw a second drop from the bottom of the mixture, that drop would contain mostly sand. Thus the composition of the sand–water mixture is not uniform—the substances have not blended smoothly and the two substances of the mixture (sand on the bottom and water on the top) remain distinct. In another example, fresh-squeezed orange juice is a mixture of juice and pulp. The pulp component floats on top of the juice component. Is your favorite pizza a mixture? The answer is yes when you consider that the pizza is a combination of distinct areas of dough, sauce, cheese, and toppings. We can therefore say that the existence of two or more distinct areas indicates a heterogeneous mixture.

A **homogeneous mixture** has constant composition throughout; it always has a single phase. Let's examine the salt–water mixture using the eyedropper. A drop of the mixture from the top of the beaker has the same composition as a drop from the bottom of the beaker. In fact, every drop of the mixture contains the same relative amounts of salt and water.

Homogeneous mixtures are also referred to as **solutions**. You are probably most familiar with solutions in a liquid form, such as cough suppressant medicine and lemonade, but solutions may contain solids, liquids, or gases. **Table 3-3** lists the various types of solution systems and gives an example of each. Solutions are very important in chemistry, and, in fact, this textbook devotes an entire chapter to the study of solutions.

The solid–solid solution known as steel is called an alloy. An alloy is a homogeneous mixture of metals, or a mixture of a metal and a nonmetal in which the metal substance is the major component. The U.S. Mint's golden dollar coin, shown in **Figure 3-13**, uses a metal alloy composed of 77% copper, 12% zinc, 7% manganese, and 4% nickel surrounding a copper core. Alloys are also used in spacecraft and automobiles. What might be the benefit of using alloys for these applications? Manufacturers combine the properties of various metals in an alloy to achieve greater strength and durability of their products.

Figure 3-13

Coins issued by the U.S. Mint are metal alloys. The combination of multiple metals gives the coins specific properties such as color, weight, and durability.

Table 3-3

Types of Solution Systems	
System	**Example**
Gas–gas	Air is primarily a mixture of nitrogen, oxygen, and argon gases.
Gas–liquid	Carbonated beverages contain carbon dioxide gas in solution.
Liquid–gas	Moist air contains water droplets in air (which is a mixture of gases).
Liquid–liquid	Vinegar contains acetic acid in water.
Solid–liquid	Sweetened powder drink contains sugar and other solid ingredients in water.
Solid–solid	Steel is an alloy of iron containing carbon.

Figure 3-14

The physical properties of the iron filings on the plate allow them to be easily separated from the sand using a magnet.

Separating Mixtures

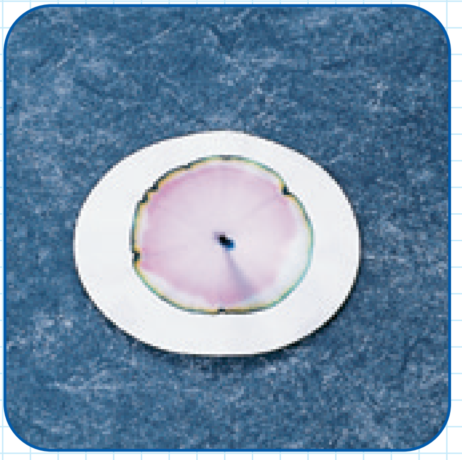
Most matter exists naturally as mixtures. For students and scientists to gain a thorough understanding of matter, it is very important to be able to do the reverse of mixing, that is, to separate mixtures into their component substances. Because the substances in a mixture are physically combined, the processes used to separate a mixture are physical processes that are based on the difference in physical properties of the substances. Sometimes it is very easy to separate a mixture; separating a mixture of pennies and nickels is not a difficult task. More difficult would be separating a mixture of sand and iron filings. Or would it be? The demonstration illustrated in **Figure 3-14** shows how the sand–iron mixture is easily separated on the basis of the unique physical properties of the substances involved. Numerous techniques have been developed that take advantage of different physical properties in order to separate mixtures.

Heterogeneous mixtures composed of solids and liquids are easily separated by filtration. **Filtration** is a technique that uses a porous barrier to separate a solid from a liquid. As **Figure 3-15** shows, the mixture is poured through a piece of filter paper that has been folded into a cone shape. The liquid passes through, leaving the solids trapped in the filter paper.

miniLAB

Separating Ink Dyes

Applying Concepts Chromatography is an important diagnostic tool for chemists. Many types of substances can be separated and analyzed using this technique. In this experiment, you will use paper chromatography to separate the dyes in water-soluble black ink.



Materials 9-oz wide-mouth plastic cups (2); round filter paper; ¼ piece of 11-cm round filter paper; scissors; pointed object, approximately 3–4 mm diameter; water-soluble black felt pen or marker

Procedure

1. Fill one of the wide-mouth plastic cups with water to about 2 cm from the top. Wipe off any water drops on the lip of the cup.
2. Place the round filter paper on a clean, dry surface. Make a concentrated ink spot in the center of the paper by firmly pressing the tip of the pen or marker onto the paper.
3. Use a sharp object to create a small hole, approximately 3–4 mm or about the diameter of a pen tip, in the center of the ink spot.
4. Roll the 1/4 piece of filter paper into a tight cone. This will act as a wick to draw the ink. Work the pointed end of the wick into the hole in the center of the round filter paper.
5. Place the paper/wick apparatus on top of the cup of water, with the wick in the water. The water will move up the wick and outward through the round paper.
6. When the water has moved to within about 1 cm of the edge of the paper (about 20 minutes), carefully remove the paper from the water-filled cup and put it on the empty cup.

Analysis

1. Make a drawing of the round filter paper and label the color bands. How many distinct dyes can you identify?
2. Why do you see different colors at different locations on the filter paper?
3. How does your chromatogram compare with those of your classmates who used other types of black felt pens or markers? Explain the differences.

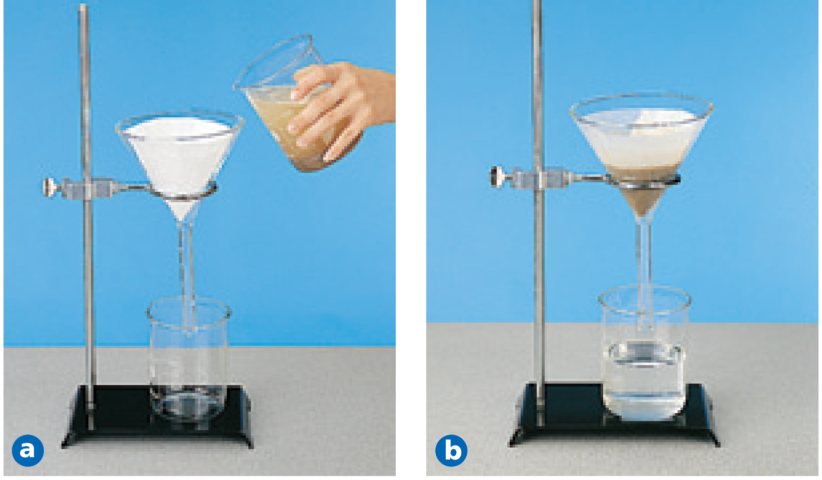


Figure 3-15

Filtration is a common technique used to remove impurities from drinking water. Clean water passes through the porous filter (a), leaving behind the impurities that can be easily discarded (b).

Most homogeneous mixtures can be separated by distillation. **Distillation** is a separation technique that is based on differences in the boiling points of the substances involved. In distillation, a mixture is heated until the substance with the lowest boiling point boils to a vapor that can then be condensed into a liquid and collected. When precisely controlled, distillation can separate substances having boiling points that differ by only a few degrees.

Did you ever make rock candy as a child? Making rock candy from a sugar solution is an example of separation by crystallization. **Crystallization** is a separation technique that results in the formation of pure solid particles of a substance from a solution containing the dissolved substance. When the solution contains as much dissolved substance as it can possibly hold, the addition of a tiny amount more often causes the dissolved substance to come out of solution and collect as crystals on some available surface. In the rock candy example, as water evaporates from the sugar–water solution, the sugar is left behind as a solid crystal on the string. Crystallization produces highly pure solids.

Chromatography is a technique that separates the components of a mixture (called the mobile phase) on the basis of the tendency of each to travel or be drawn across the surface of another material (called the stationary phase). The **miniLAB** in this section describes how you can separate a solution such as ink into its components as it spreads across a stationary piece of paper. The separation occurs because the various components of the ink spread through the paper at different rates.

Section 3.3 Assessment

15. How do mixtures and substances differ?

16. Consider a mixture of water, sand, and oil. How many phases are present? How could you separate this mixture into individual substances?

17. Classify each of the following as either a heterogeneous or homogeneous mixture.

a. orange juice

b. tap water

c. steel (a blend of iron and carbon)

d. air

e. raisin muffin

18. **Thinking Critically** When 50 mL of ethanol is mixed with 50 mL of water, a solution forms. The volume of the final solution is less than 100 mL. Propose an explanation for this phenomenon. (*Hint*: Consider what you know about the space between particles in liquids.)

19. **Applying Concepts** Describe the separation technique that could be used to separate each of the following mixtures.

a. two colorless liquids

b. a nondissolving solid mixed with a liquid

c. red and blue marbles of same size and mass

Section 3.4

Elements and Compounds

Objectives

- **Distinguish** between elements and compounds.
- **Describe** the organization of elements on the periodic table.
- **Explain** how all compounds obey the laws of definite and multiple proportions.

Vocabulary

element
periodic table
compound
law of definite proportions
percent by mass
law of multiple proportions

To this point you've examined many of the properties of matter. You've also learned how scientists have organized, classified, and described matter by arranging it into various subcategories of components. But there remains another fundamental level of classification of matter: the classification of pure substances as elements or compounds.

Elements

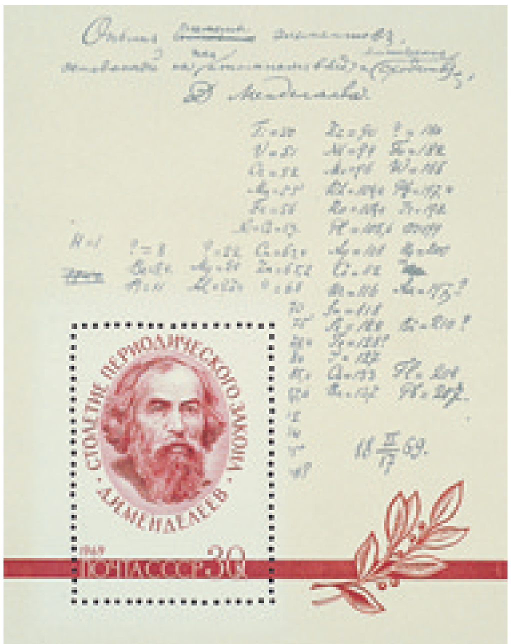
Recall that earlier in this chapter you considered the diversity of your surroundings in terms of matter. Although the diversity is astounding, in reality all matter can be broken down into a relatively small number of basic building blocks called elements. An **element** is a pure substance that cannot be separated into simpler substances by physical or chemical means. On Earth, 91 elements occur naturally. Copper, oxygen, and gold are examples of naturally occurring elements. There are also several elements that do not exist naturally but have been developed by scientists.

Each element has a unique chemical name and symbol. The chemical symbol consists of one, two, or three letters; the first letter is always capitalized and the remaining letter(s) are always lowercase. Why has so much effort been given to naming the elements? The names and symbols of the elements are universally accepted by scientists in order to make the communication of chemical information possible.

The 91 naturally occurring elements are not equally abundant. For example, hydrogen is estimated to make up approximately 75% of the mass of the universe. Oxygen and silicon together comprise almost 75% of the mass of Earth's crust, while oxygen, carbon, and hydrogen account for more than 90% of the human body. Francium, on the other hand, is one of the least abundant naturally occurring elements. It is estimated that there is probably less than 20 grams of francium dispersed throughout Earth's crust. To put that into perspective, the total mass of francium is approximately equal to the mass of your pencil or pen.

Figure 3-16

Although many early scientists have contributed to the modern organization of the elements, Mendeleev's system of rows and columns was a revolutionary advancement.



A first look at the periodic table As many new elements were being discovered in the early nineteenth century, chemists began to see patterns of similarities in the chemical and physical properties of particular sets of elements. Several schemes for organizing the elements on the basis of these similarities were proposed, with varying degrees of success. In 1869, the Russian chemist Dmitri Mendeleev made a significant contribution to the effort. Mendeleev devised the chart shown in **Figure 3-16**, which organized all of the elements that were known at the time into rows and columns based on their similarities and their masses. Mendeleev's organizational table was the first version of what has been further developed into the **periodic table** of elements. The periodic table organizes the elements into a grid of horizontal rows called periods and vertical columns called groups or families. Elements in the same group have similar chemical and physical properties. The table is called "periodic" because the pattern of similar properties repeats as you move from period to period.

One of the brilliant aspects of Mendeleev's original table was that its structure could accommodate elements that did not even exist at

the time. Notice the blank spots in Mendeleev's table. By analyzing the similarities among the elements and their pattern of repetition, Mendeleev was able to predict the properties of elements that were yet to be discovered.

In most cases, Mendeleev's predictions (and the blanks in the table) closely matched the characteristics of new elements as they were discovered. **Figure 3-18** on pages 72–73 shows samples of the elements in their arrangement in the periodic table. The standard modern version of the periodic table includes more than 100 elements. You'll study the periodic table in greater detail later in this textbook. In fact, the periodic table remains a dynamic tool as scientists continue to discover new elements.

Compounds

Take a moment to recall what you have learned about the organization of matter, using **Figure 3-17** as a guide. You know that matter is classified as pure substances and mixtures. As you learned in the previous section, mixtures can be homogeneous or heterogeneous. You also know that elements are pure substances that cannot be separated into simpler substances. There is yet another classification of pure substances—compounds. A **compound** is a combination of two or more different elements that are combined chemically. Most of the substances that you are familiar with and, in fact, much of the matter of the universe are compounds. Water, table salt, table sugar, and aspirin are examples of common compounds.

Today, there are approximately 10 million known compounds, and new compounds continue to be developed and discovered at the rate of about 100 000 per year. Can you recall some of the medicinal compounds that have made headlines in recent years? There appears to be no limit to the number of compounds that can be made or that will be discovered. Considering this virtually limitless potential, several organizations have assumed the task of collecting data and indexing the known chemical compounds. These organizations maintain huge databases that allow researchers to access information on existing compounds. The databases and retrieval tools enable scientists to build the body of chemical knowledge in an efficient manner.

The chemical symbols of the periodic table make it easy to write the formulas for chemical compounds. For example, table salt, or sodium chloride, is composed of one part sodium (Na) and one part chlorine (Cl), and its chemical formula is NaCl. Water is composed of two parts hydrogen (H) to one part oxygen (O), and its formula is H_2O.

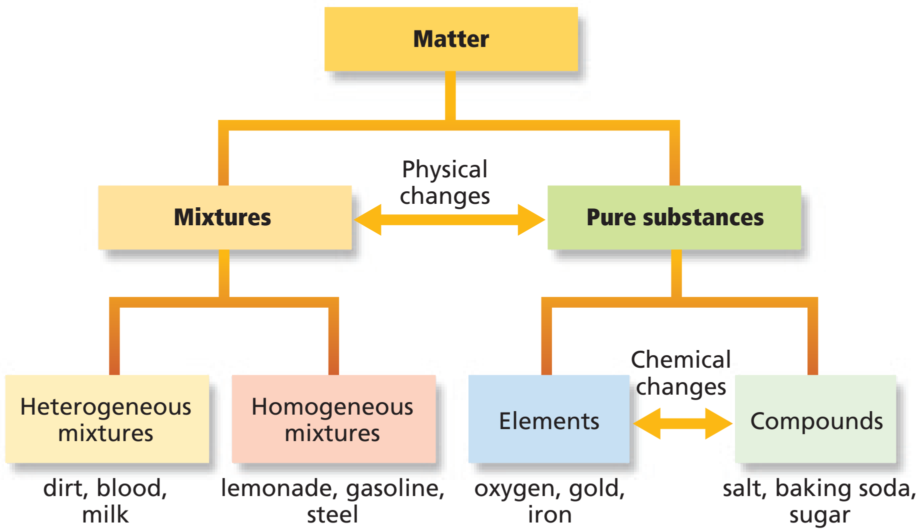


Figure 3-17

The concept of matter is far-reaching and can be overwhelming. But, when broken down as shown here, it becomes clear how elements, compounds, substances, and mixtures define all matter.

Periodic Table

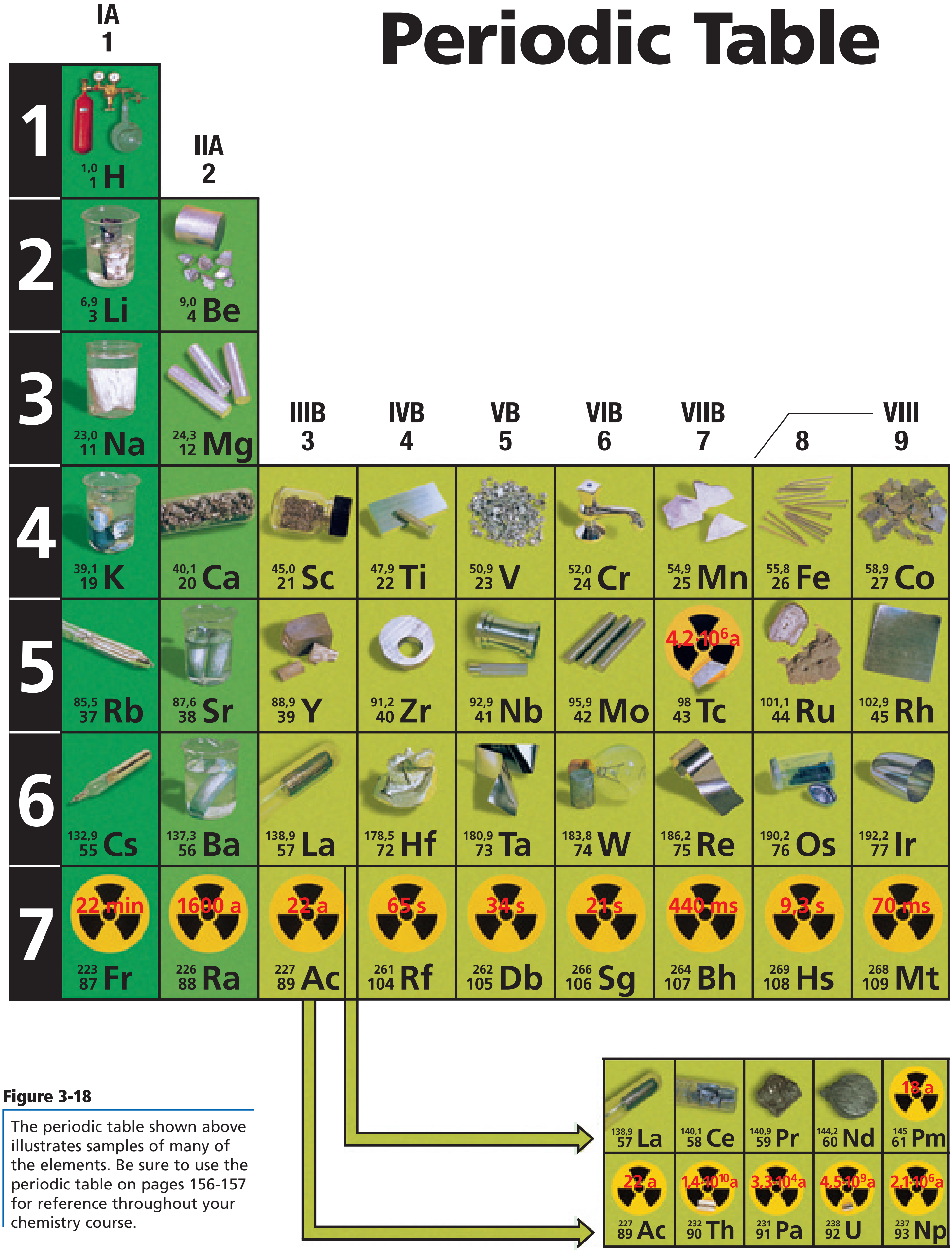


Figure 3-18

The periodic table shown above illustrates samples of many of the elements. Be sure to use the periodic table on pages 156-157 for reference throughout your chemistry course.

of the Elements

10	IB 11	IIB 12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18
								4,0 2 He
			10,8 5 B	12,0 6 C	14,0 7 N	16,0 8 O	19,0 9 F	20,2 10 Ne
			27,0 13 Al	28,1 14 Si	31,0 15 P	32,1 16 S	35,5 17 Cl	39,9 18 Ar
58,7 28 Ni	63,5 29 Cu	65,4 30 Zn	69,7 31 Ga	72,6 32 Ge	74,9 33 As	79,0 34 Se	79,9 35 Br	83,8 36 Kr
106,4 46 Pd	107,9 47 Ag	112,4 48 Cd	114,8 49 In	118,7 50 Sn	121,8 51 Sb	127,6 52 Te	126,9 53 I	131,3 54 Xe
195,1 78 Pt	197,0 79 Au	200,6 80 Hg	204,4 81 Tl	207,2 82 Pb	209,0 83 Bi	102 a 209 84 Po	8,1 h 210 85 At	3,8 d 222 86 Rn
118 ms 273 110	1,5 ms 272 111	0,24 ms 277 112						

150,4 62 Sm	152,0 63 Eu	157,3 64 Gd	158,9 65 Tb	162,5 66 Dy	164,9 67 Ho	167,3 68 Er	168,9 69 Tm	173,0 70 Yb	175,0 71 Lu
$8{,}0 \cdot 10^7$ a 244 94 Pu	7400 a 243 95 Am	$1{,}6 \cdot 10^7$ a 247 96 Cm	1400 a 247 97 Bk	900 a 251 98 Cf	472 d 252 99 Es	101 d 257 100 Fm	52 d 258 101 Md	58 min 259 102 No	3,6 h 262 103 Lr

Figure 3-19

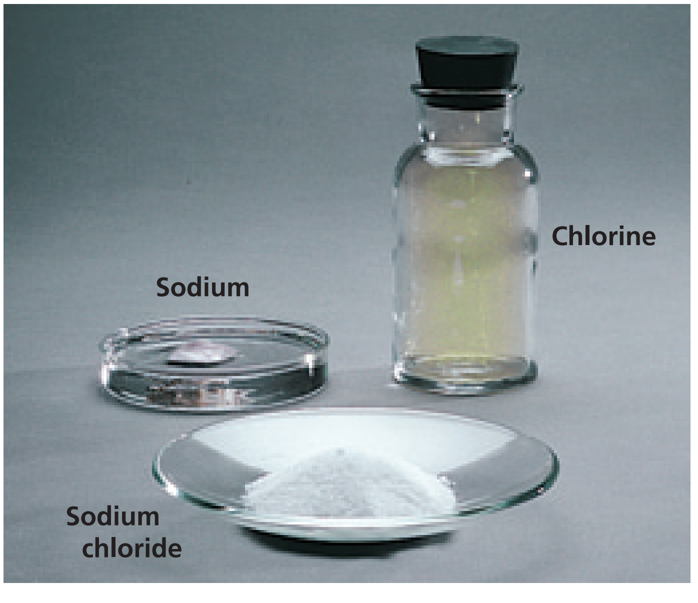
This classic apparatus, called a Hoffman apparatus, and other similar designs are used to separate water into its components.

Unlike elements, compounds can be broken down into simpler substances by chemical means. In general, compounds that naturally occur are more stable than the individual component elements. To separate a compound into its elements often requires external energy such as heat or electricity. **Figure 3-19** shows the apparatus used to produce the chemical change of water into its component elements of hydrogen and oxygen through a process called electrolysis. Here, one end of a long platinum electrode is exposed to the water in the tube and the other end is attached to a power source. An electric current splits water into hydrogen gas in the compartment on the right and oxygen gas in the compartment on the left. Because water is composed of two parts hydrogen and one part oxygen, there is twice as much hydrogen gas than oxygen gas.

The properties of a compound are different from those of its component elements. The example of water in **Figure 3-19** illustrates this fact. Water is a stable compound that is liquid at room temperature. When water is broken down into its components, it is obvious that hydrogen and oxygen are dramatically different than the liquid they form when combined. Oxygen and hydrogen are tasteless, odorless gases that vigorously undergo chemical reactions with many elements. This difference in properties is a result of a chemical reaction between the elements. **Figure 3-20** shows the component elements (sodium and chlorine) of the compound commonly called table salt (sodium chloride). When sodium and chlorine react with each other, the compound sodium chloride is formed. Note how different the properties of sodium chloride are from its component elements. Sodium is a highly reactive element that fizzes and burns when added to water. Chlorine is a poisonous, pale green gas. Sodium chloride, however, is a white, unreactive solid that flavors many of the foods you eat.

Figure 3-20

Compounds such as sodium chloride (table salt) are often remarkably different from the components that comprise them.



Law of Definite Proportions

An important characteristic of compounds is that the elements comprising them combine in definite proportions by mass. This observation is so fundamental that it is summarized as the **law of definite proportions**. This law states that, regardless of the amount, a compound is always composed of the same elements in the same proportion by mass. For example, consider the compound table sugar (sucrose), which is composed of carbon, hydrogen, and oxygen. The analysis of 20.00 g of sucrose from a bag of sugar is given in **Table 3-4**. Note that in Column 1 the sum of the individual masses of the elements equals 20.00 g, the amount of sucrose that was analyzed. This demonstrates the law of conservation of mass as applied to compounds: The mass of the compound is equal to the sum of the masses of the elements that make up the compound. Column 2 shows the ratio of the mass of each element to the total mass of the compound as a percentage called the **percent by mass**.

$$\text{percent by mass (\%)} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

Table 3-4

Sucrose Analysis from Bag Sugar		
	Column 1	Column 2
Element	**Analysis by mass (g)**	**Percent by mass (%)**
Carbon	8.44 g carbon	$\frac{8.44 \text{ g C}}{20.00 \text{ g sucrose}} \times 100 = 42.2\%$ carbon
Hydrogen	1.30 g hydrogen	$\frac{1.30 \text{ g H}}{20.00 \text{ g sucrose}} \times 100 = 6.50\%$ hydrogen
Oxygen	10.26 g oxygen	$\frac{10.26 \text{ g O}}{20.00 \text{ g sucrose}} \times 100 = 51.30\%$ oxygen
Total	20.00 g sucrose	= 100.0%

Now let's suppose you analyzed 500.0 g of sucrose isolated from a sample of sugar cane. The analysis is shown in **Table 3-5**. Note in Column 2 that the percent by mass values equal those in Column 2 in **Table 3-4.** Compounds with the same mass proportions must be the same compound; conversely, compounds with different mass proportions must be different compounds. Thus, you can conclude that, although the two samples of sucrose are from different sources, they have the same composition and they must be the same compound.

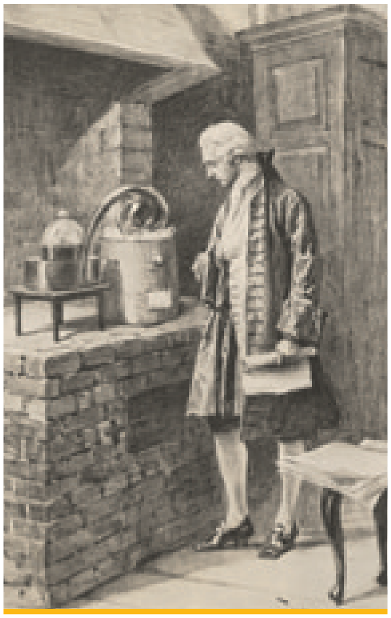
Table 3-5

Sucrose Analysis from Sugar Cane		
	Column 1	Column 2
Element	**Analysis by mass (g)**	**Percent by mass (%)**
Carbon	211.0 g carbon	$\frac{211.0 \text{ g C}}{500.0 \text{ g sucrose}} \times 100 = 42.20\%$ carbon
Hydrogen	32.5 g hydrogen	$\frac{32.5 \text{ g H}}{500.0 \text{ g sucrose}} \times 100 = 6.50\%$ hydrogen
Oxygen	256.5 g oxygen	$\frac{256.5 \text{ g O}}{500.0 \text{ g sucrose}} \times 100 = 51.30\%$ oxygen
Total	500.0 g sucrose	= 100.00%

History CONNECTION

Antoine-Laurent Lavoisier (1743–1794) is recognized as the father of modern chemistry. While his fellow scientists tried to explain matter based on the elements fire, earth, air, and water, Lavoisier performed some of the first quantitative chemical experiments. His data and observations led to the statement of the law of conservation of mass. He also studied the nature of combustion and devised a system of naming elements.

Lavoisier held many public offices in France in which he attempted to reform the French monetary and taxation system and farming methods. He also supervised the French government's manufacture of gunpowder. During the Reign of Terror that followed the French Revolution, Lavoisier and other members of Fermés Generalé were arrested, tried, and condemned to the guillotine.



Practice!

For more practice with percent by mass and law of definite proportions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

20. A 78.0-g sample of an unknown compound contains 12.4 g of hydrogen. What is the percent by mass of hydrogen in the compound?

21. If 1.0 g of hydrogen reacts completely with 19.0 g of fluorine, what is the percent by mass of hydrogen in the compound that is formed?

22. If 3.5 g of X reacts with 10.5 g of Y to form the compound XY, what is the percent by mass of X in the compound? The percent by mass of Y?

23. Two unknown compounds are tested. Compound I contains 15.0 g of hydrogen and 120.0 g of oxygen. Compound II contains 2.0 g of hydrogen and 32.0 g of oxygen. Are the compounds the same?

24. All you know about two unknown compounds is that they have the same percent by mass of carbon. With only this information, can you be sure the two compounds are the same?

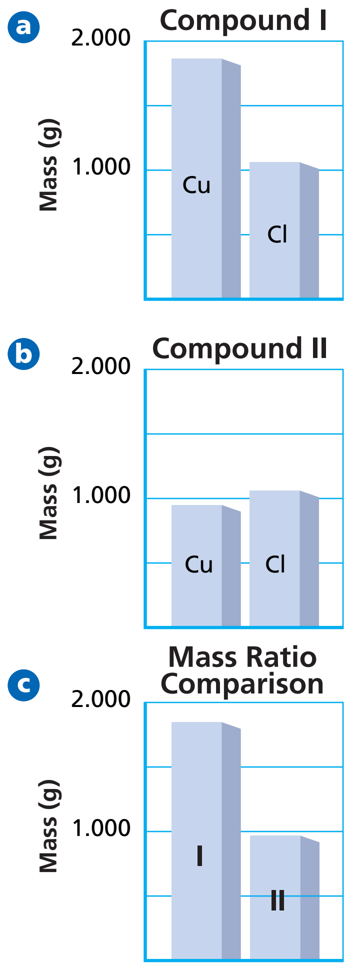


Figure 3-21

Bar graph (a) compares the relative masses of copper and chlorine in Compound I and bar graph (b) compares the relative masses of copper and chlorine in Compound II. (c) A comparison between the relative masses of copper in both compounds shows a 2:1 ratio.

Law of Multiple Proportions

Compounds composed of different elements are obviously different compounds. Can compounds that are composed of the same elements differ from each other? The answer is yes because those different compounds have different mass compositions. The **law of multiple proportions** states that when different compounds are formed by a combination of the same elements, different masses of one element combine with the same relative mass of the other element in a ratio of small whole numbers. Ratios compare the relative amounts of any items or substances. The comparison can be expressed using numbers separated by a colon or as a fraction. With regard to the law of multiple proportions, ratios express the relationship of elements in a compound.

The two distinct compounds water (H_2O) and hydrogen peroxide (H_2O_2) illustrate the law of multiple proportions. Each compound contains the same elements (hydrogen and oxygen). Water is composed of two parts hydrogen (the element that is present in the same amount in both compounds) to one part oxygen (the element that is present in different amounts in both compounds). Hydrogen peroxide is composed of two parts hydrogen and two parts oxygen. Hydrogen peroxide differs from water in that it has twice as much oxygen. When we compare the mass of oxygen in hydrogen peroxide to the mass of oxygen in water, we get the ratio 2:1.

In another example, copper (Cu) reacts with chlorine (Cl) under different sets of conditions to form two different compounds. **Table 3-6** provides an analysis of their composition. Note that the two copper compounds must be different because they have different percents by mass. Compound I contains 64.20% copper; compound II contains 47.27% copper. Compound I contains 35.80% chlorine; compound II contains 52.73% chlorine.

Table 3-6

Analysis Data of Two Copper Compounds

Compound	% Cu	% Cl	Mass copper (g) in 100.0 g of compound	Mass chlorine (g) in 100.0 g of compound	Mass ratio $\left(\frac{\text{mass Cu}}{\text{mass Cl}}\right)$
I	64.20	35.80	64.20	35.80	1.793 g Cu/1 g Cl
II	47.27	52.73	47.27	52.73	0.8964 g Cu/1 g Cl

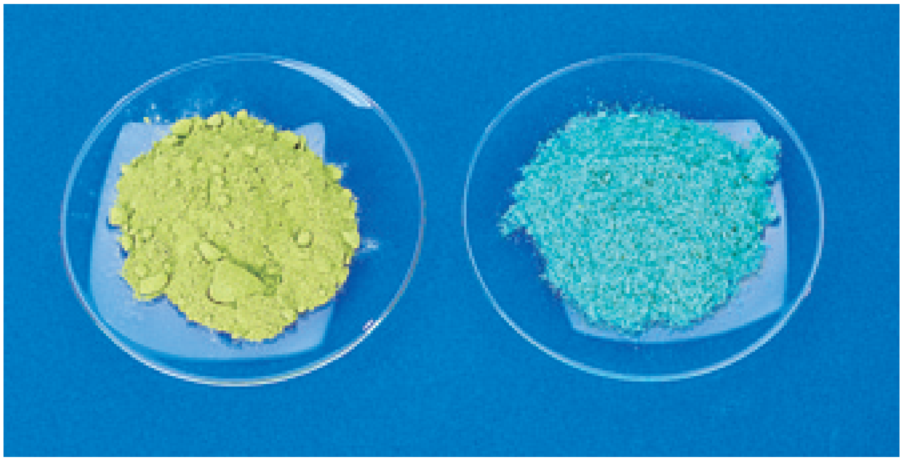


Figure 3-22

Analyses of the mass ratios of the two copper chloride compounds shown here indicate that they are indeed different compounds. The calculated mass ratio of compound I to compound II is 2.000 and fits the definition of the law of multiple proportions.

Compare the ratio of the mass of copper to the mass of chlorine for each compound (see the last column of **Table 3-6** and **Figure 3-21**). You'll notice that the mass ratio of copper to chlorine in compound I (1.793) is two times the mass ratio of copper to chlorine in compound II (0.896).

$$\frac{\text{mass ratio compound I}}{\text{mass ratio compound II}} = \frac{1.793 \text{ g Cu/g Cl}}{0.8964 \text{ g Cu/g Cl}} = 2.000$$

As the law of multiple proportions states, the different masses of copper that combine with a fixed mass of chlorine in the two different copper compounds, shown in **Figure 3-22,** can be expressed as a small whole-number ratio, in this case 2:1.

Considering that there is a finite number of elements that exist today and an exponentially greater number of compounds that are composed of these elements under various conditions, it becomes clear how important the law of multiple proportions is in chemistry.

Section 3.4 Assessment

25. How are elements and compounds similar? How are they different?

26. What is the basic organizing feature of the periodic table of elements?

27. Explain how the law of definite proportions applies to compounds.

28. What type of compounds are compared in the law of multiple proportions?

29. **Thinking Critically** Name two elements that have properties similar to those of element potassium (K). To those of krypton (Kr).

30. **Interpreting Data** Complete the following table and then analyze the data to determine if compounds I and II are the same compound. If the compounds are different, use the law of multiple proportions to show the relationship between them.

Analysis Data of Two Iron Compounds

Compound	Total mass (g)	Mass Fe (g)	Mass O (g)	Mass % Fe	Mass % O
I	75.00	52.46	22.54		
II	56.00	43.53	12.47		

CHEMLAB 3

Matter and Chemical Reactions

One of the most interesting characteristics of matter, and one that drives the study and exploration of chemistry, is the fact that matter changes. By examining a dramatic chemical reaction, such as the reaction of the element copper and the compound silver nitrate in a water solution, you can readily observe chemical change. Drawing on one of the fundamental laboratory techniques introduced in this chapter, you can separate the products. Then, you will use a flame test to confirm the identity of the products.

Problem

Is there evidence of a chemical reaction between copper and silver nitrate? If so, which elements reacted and what is the name of the compound they formed?

Objectives

- **Observe** the reactants as they change into product.
- **Separate** a mixture by filtration.
- **Predict** the names of the products.

Materials

copper wire
$AgNO_3$ solution
sandpaper
stirring rod
50-mL graduated cylinder
50-mL beaker
funnel
filter paper
250-mL Erlenmeyer flask
ring stand
small iron ring
plastic petri dish
paper clip
Bunsen burner
tongs

Safety Precautions

- **Always wear safety goggles, gloves, and lab apron.**
- **Silver nitrate is toxic and will harm skin and clothing.**
- **Use caution around a flame.**

Pre-Lab

1. Read the entire **CHEMLAB**.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your observations.
3. Define the terms physical property and chemical property. Give an example of each.
4. Form a hypothesis regarding what you might observe if
   a. a chemical change occurs.
   b. a physical change occurs.
5. Distinguish between a homogeneous mixture and a heterogeneous mixture.

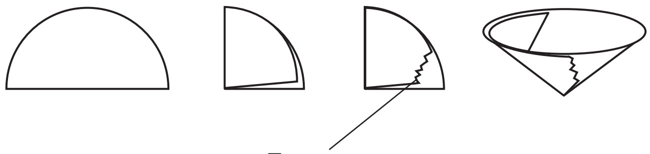
Reaction Observations

Time (min)	Observations

Procedure

1. Obtain 8 cm of copper wire. Rub the copper wire with the sandpaper until it is shiny.
2. Measure approximately 25 mL $AgNO_3$ (silver nitrate) solution into a 50-mL beaker. **CAUTION:** *Do not allow to contact skin or clothing.*
3. Make and record an observation of the physical properties of the copper wire and $AgNO_3$ solution.

4. Coil the piece of copper wire to a length that will fit into the beaker. Make a hook on the end of the coil to allow the coil to be suspended from the stirring rod.
5. Hook the coil onto the middle of the stirring rod. Place the stirring rod across the top of the beaker immersing some of the coil in the $AgNO_3$ solution.
6. Make and record observations of the wire and the solution every five minutes for 20 minutes.
7. Use the ring stand, small iron ring, funnel, Erlenmeyer flask, and filter paper to set up a filtration apparatus. Attach the iron ring to the ring stand. Adjust the height of the ring so the end of the funnel is inside the neck of the Erlenmeyer flask.
8. To fold the filter paper, examine the diagram below. Begin by folding the circle in half, then fold in half again. Tear off the lower right corner of the flap that is facing you. This will help the filter paper stick better to the funnel. Open the folded paper into a cone. Place the filter paper cone in the funnel.



9. Remove the coil from the beaker and dispose of it as directed by your teacher. Some of the solid product may form a mixture with the liquid in the beaker. Decant the liquid by slowly pouring it down the stirring rod into the funnel. Solid product will be caught in the filter paper. Collect the filtrate—the liquid that runs through the filter paper—in the Erlenmeyer flask.
10. Transfer the clear filtrate to a petri dish.
11. Adjust a Bunsen burner flame until it is blue. Hold the paper clip with tongs in the flame until no additional color is observed. **CAUTION:** *The paper clip will be very hot.*
12. Using tongs, dip the hot paper clip in the filtrate. Then, hold the paper clip in the flame. Record the color you observe.

Cleanup and Disposal

1. Dispose of materials as directed by your teacher.
2. Clean and return all lab equipment to its proper place.
3. Wash hands thoroughly.

Analyze and Conclude

1. **Classifying** Which type of mixture is silver nitrate in water? Which type of mixture is formed in step 6? Explain.
2. **Observing and Inferring** Describe the changes you observed in step 6. Is there evidence a chemical change occurred? Why?
3. **Predicting** Predict the products formed in step 6. You may not know the exact chemical name, but you should be able to make an intuitive prediction.
4. **Using Resources** Use resources such as the *CRC Handbook of Chemistry and Physics*, the *Merck Index,* or the Internet to determine the colors of silver metal and copper nitrate in water. Compare this information with your observations of the reactants and products in step 6.
5. **Identifying** Metals emit characteristic colors in flame tests. Copper emits blue-green light. Do your observations in step 12 confirm the presence of copper in the filtrate collected in step 9?
6. **Communicating** Express in words the chemical equation that represents the reaction that occurred in step 6.
7. **Error Analysis** Compare your recorded observations with those of several other lab teams. Explain any differences.

Real-World Chemistry

1. Analytical chemists determine the chemical composition of matter. Two major branches of analytical chemistry are qualitative analysis—determining what is in a substance—and quantitative analysis—measuring how much substance. Research and report on a career as an analytical chemist in the food industry.

CHEMISTRY and Society

Green Buildings

Until the Industrial Revolution, the amount of carbon dioxide (CO_2) in the atmosphere was fairly constant. Since the Industrial Revolution, however, the burning of fossil fuels has contributed to a significant increase in the amount of carbon dioxide in the atmosphere. As the level of carbon dioxide increases, Earth gradually warms up. Too much CO_2 in the atmosphere can change the conditions on Earth.

Another major source of carbon dioxide may be in the foundation of your building or on the sidewalks near your school. The production of cement, the key ingredient in concrete, releases tremendous amounts of carbon dioxide into the atmosphere. Chemistry may allow engineers to build "green buildings," that are still practical yet have less of an impact on the environment.

Producing Cement

Cement generally begins with a mixture of limestone and sand placed in a kiln, which heats it to about 1480°C. As the mixture is heated, its chemical and physical properties change. After heating, the solid that remains is ground into a fine powder. This is cement. To make concrete, the cement is mixed with fine particles, such as sand, coarse particles, such as crushed stone, and water.

During the production of cement, carbon dioxide is released in two ways. First, when the limestone is heated it changes into lime and carbon dioxide. Second, the electrical energy used to heat the kiln is usually supplied by a power plant that burns fossil fuels, such as coal. Fossil fuels also release carbon dioxide and other substances.

Using Flyash

One way to reduce the amount of carbon dioxide released into the atmosphere is to find a replacement for cement in concrete. One such replacement is a substance known as flyash. Flyash is a waste product that accumulates in the smokestacks of power plants when ground coal is burned. It is a fine gray powder that consists of tiny glass beads.

Using flyash offers several advantages. Flyash ordinarily is dumped in landfills. Replacing cement with flyash can reduce CO_2 emissions and prevent

tons of waste from piling up in landfills. Flyash also produces better concrete. Traditional concrete has weak zones where tiny cracks allow water to flow through. Flyash contains fine particles that fill spaces and keep moisture out. Flyash also protects the steel surrounding the concrete, makes the concrete easier to work with, and extends the life of the concrete structure. In fact, flyash is so reliable the Romans used natural materials similar to flyash to build the concrete dome of the Pantheon.

Solutions to environmental problems require a willing commitment from scientists, architects, builders, and owners to look for ways to build durable structures and protect the environment.

Investigating the Issue

1. **Communicating Ideas** Write a pamphlet for people who are building new homes telling them about the importance of green buildings.
2. **Using the Internet** Investigate issues that influence the decision to use flyash. Discuss the advantages and disadvantages of flyash.

CHEMISTRY Online

CLICK HERE

Visit the Chemistry Web site at **science.glencoe.com** to find links to more information about flyash and green buildings.

Summary

3.1 Properties of Matter

- A substance is a form of matter with a uniform and unchanging composition.
- Physical properties can be observed without altering a substance's composition. Chemical properties describe a substance's ability to combine with or change into one or more new substances.
- Both physical and chemical properties are affected by external conditions such as temperature and pressure.
- The three common states of matter are solid, liquid, and gas.

3.2 Changes in Matter

- A physical change alters the physical properties of a substance without changing its composition.
- A chemical change, also known as a chemical reaction, involves a change in a substance's composition.
- In a chemical reaction, reactants form products.
- The law of conservation of mass states that mass is neither created nor destroyed during a chemical reaction; it is conserved.

3.3 Mixtures of Matter

- A mixture is a physical blend of two or more pure substances in any proportion.
- Solutions are homogeneous mixtures.
- Mixtures can be separated by physical means. Common separation techniques include filtration, distillation, crystallization, and chromatography.

3.4 Elements and Compounds

- Elements are substances that cannot be broken down into simpler substances by chemical or physical means.
- The elements are organized in the periodic table of elements.
- A compound is a chemical combination of two or more elements. Properties of compounds differ from the properties of their component elements.
- The law of definite proportions states that a compound is always composed of the same elements in the same proportions.
- The law of multiple proportions states that if elements form more than one compound, those compounds will have compositions that are small, whole-number multiples of each other.

Key Equations and Relationships

- law of conservation of mass (p. 63)
  $\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$
- percent by mass $= \dfrac{\text{Mass}_{\text{element}}}{\text{Mass}_{\text{compound}}} \times 100$
  (p. 75)

Vocabulary

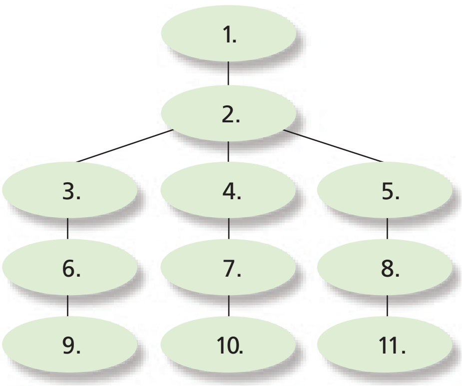
- chemical change (p. 62)
- chemical property (p. 57)
- chromatography (p. 69)
- compound (p. 71)
- crystallization (p. 69)
- distillation (p. 69)
- element (p. 70)
- extensive properties (p. 56)
- filtration (p. 68)
- gas (p. 59)
- heterogeneous mixture (p. 67)
- homogeneous mixture (p. 67)
- intensive properties (p. 56)
- law of conservation of mass (p. 63)
- law of definite proportions (p. 75)
- law of multiple proportions (p. 76)
- liquid (p. 58)
- mixture (p. 66)
- percent by mass (p. 75)
- periodic table (p. 70)
- physical changes (p. 61)
- physical property (p. 56)
- solid (p. 58)
- solution (p. 67)
- states of matter (p. 58)
- substance (p. 55)
- vapor (p. 59)

CHEMISTRY Online CLICK HERE

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 3 Assessment.*

Concept Mapping

31. Organize the following terms into a logical concept map: state, physical properties, virtually incompressible, solid, gas, liquid, tightly packed particles, compressible, incompressible, particles far apart, loosely packed particles.



Mastering Concepts

32. List three examples of substances. Explain why each is a substance. (3.1)

33. List at least three physical properties of tap water. (3.1)

34. Identify each of the following as an extensive or intensive physical property. (3.1)

- **a.** melting point
- **b.** mass
- **c.** density
- **d.** length

35. "Properties are not affected by changes in temperature and pressure." Is this statement true or false? Explain. (3.1)

36. Classify each of the following as either solid, liquid, or gas at room temperature. (3.1)

- **a.** milk
- **b.** air
- **c.** copper
- **d.** helium
- **e.** diamond
- **f.** candle wax

37. Classify each of the following as a physical property or a chemical property. (3.1)

- **a.** aluminum has a silvery color
- **b.** gold has a density of 19 g/cm^3
- **c.** sodium ignites when dropped in water
- **d.** water boils at 100°C
- **e.** silver tarnishes
- **f.** mercury is a liquid at room temperature

38. A carton of milk is poured into a bowl. Describe the changes that occur in the milk's shape and volume. (3.1)

39. Classify each of the following as a physical change or a chemical change. (3.2)

- **a.** breaking a pencil in two
- **b.** water freezing and forming ice
- **c.** frying an egg
- **d.** burning wood
- **e.** leaves turning color in the fall

40. Is a change in phase a physical change or a chemical change? Explain. (3.2)

41. List four indicators that a chemical change has probably taken place. (3.2)

42. Iron and oxygen combine to form iron oxide (rust). List the reactants and products of this reaction. (3.2)

43. Use **Table 3-1** to identify a substance that undergoes a phase change as its temperature increases from −250°C to −210°C. What phase change takes place? (3.2)

44. After burning for three hours, a candle has lost half of its mass. Explain why this example does not violate the law of conservation of mass. (3.2)

45. Describe the difference between a physical change and a chemical change. (3.2)

46. Describe the characteristics of a mixture. (3.3)

47. Describe a method that could be used to separate each of the following mixtures. (3.3)

- **a.** iron filings and sand
- **b.** sand and salt
- **c.** the components of ink
- **d.** helium and oxygen gases

48. "A mixture is the chemical bonding of two or more substances in any proportion." Is this statement true or false. Explain.

49. Which of the following are the same and which are different? (3.3)

a. a substance and a pure substance
b. a heterogeneous mixture and a solution
c. a substance and a mixture
d. a homogeneous mixture and a solution

50. Describe how a homogeneous mixture differs from a heterogeneous mixture. (3.3)

51. A chemistry professor has developed a laboratory task to give her students practical experience using basic separation techniques. She prepares a liquid solution of water and another compound. Assuming you are a student in the class, name the technique you would use to separate and identify the components. Give specific details of the method.

52. State the definition of an element. (3.4)

53. Name the elements contained in the following compounds. (3.4)

a. sodium chloride (NaCl) **c.** ethanol (C_2H_6O)
b. ammonia (NH_3) **d.** bromine (Br_2)

54. How many naturally occurring elements are found on Earth? Approximately how many synthetic elements have been identified? (3.4)

55. What was Dmitri Mendeleev's major contribution to the field of chemistry? (3.4)

56. Is it possible to distinguish between an element and a compound? Explain. (3.4)

57. How are the properties of a compound related to those of the elements that comprise it? (3.4)

58. How are the elements contained within a group on the periodic table related? (3.4)

59. Which law states that a compound always contains the same elements in the same proportion by mass? (3.4)

Mastering Problems

Properties of Matter (3.1)

60. A scientist is given the task of identifying an unknown compound on the basis of its physical properties. The substance is a white solid at room temperature. Attempts to determine its boiling point were unsuccessful. Using **Table 3-1**, name the unknown compound.

Conservation of Mass (3.2)

61. A 28.0-g sample of nitrogen gas combines completely with 6.0 g of hydrogen gas to form ammonia. What is the mass of ammonia formed?

62. A substance breaks down into its component elements when it is heated. If 68.0 grams of the substance is present before it is heated, what is the combined mass of the component elements after heating?

63. A 13.0-g sample of X combines with a 34.0-g sample of Y to form the compound XY_2. What is the mass of the reactants?

64. Sodium chloride can be formed by the reaction of sodium metal and chlorine gas. If 45.98 g of sodium combines with an excess of chlorine gas to form 116.89 g sodium chloride, what mass of chlorine gas is used in the reaction?

65. Copper sulfide is formed when copper and sulfur are heated together. In this reaction, 127 g of copper reacts with 41 g of sulfur. After the reaction is complete, 9 g of sulfur remains unreacted. What is the mass of copper sulfide formed?

Law of Definite Proportions (3.4)

66. A 25.3-g sample of an unknown compound contains 0.8 g of oxygen. What is the percent by mass of oxygen in the compound?

67. Magnesium combines with oxygen to form magnesium oxide. If 18.06 g of magnesium reacts completely with 6.96 g of oxygen, what is the percent by mass of oxygen in magnesium oxide?

68. When mercury oxide is heated, it decomposes into mercury and oxygen. If 28.4 g of mercury oxide decomposes, producing 2.0 g oxygen, what is the percent by mass of mercury in mercury oxide?

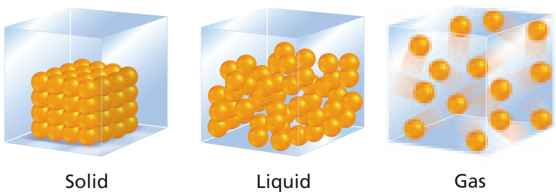
Law of Multiple Proportions (3.4)

69. Carbon reacts with oxygen to form two different compounds. Compound I contains 4.82 g carbon for every 6.44 g of oxygen. Compound II contains 20.13 g carbon for every 53.7 g of oxygen. What is the ratio of carbon to a fixed mass of oxygen for the two compounds?

Mixed Review

Sharpen your problem-solving skills by answering the following.

70. Which state of matter is the most compressible? The least? Explain why.



CHAPTER 3 ASSESSMENT

71. Classify each of the following as a homogeneous mixture or a heterogeneous mixture. (3.3)

a. brass (an alloy of zinc and copper)
b. a salad
c. blood
d. powder drink mix dissolved in water

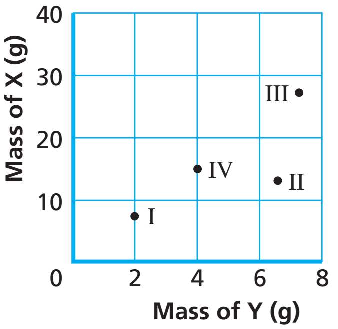
72. Phosphorus combines with hydrogen to form phosphine. In this reaction, 123.9 g of phosphorus combines with excess hydrogen to produce 129.9 g of phosphine. After the reaction, 310 g of hydrogen remains unreacted. What mass of hydrogen is used in the reaction? What was the initial mass of hydrogen before the reaction?

73. A sample of a certain lead compound contains 6.46 grams of lead for each gram of oxygen. A second sample has a mass of 68.54 g and contains 28.76 g of oxygen. Are the two samples the same?

Thinking Critically

74. Applying Concepts Air is a mixture of many gases, primarily nitrogen, oxygen, and argon. Could distillation be used to separate air into its component gases? Explain.

75. Interpreting Data A compound contains elements X and Y. Four samples with different masses were analyzed, and the masses of X and Y in each sample were plotted on a graph. The samples are labeled I, II, III, and IV.



a. Which samples are from the same compound? How do you know?
b. What is the approximate ratio of mass X to mass Y in the samples that are from the same compound?
c. What is the approximate ratio of mass X to mass Y in the sample(s) that are not from the same compound?

Writing in Chemistry

76. Select a synthetic element and prepare a short written report on its development. Be sure to cover recent discoveries, list major research centers that conduct this type of research, and describe the properties of the synthesized element.

77. Research the life of a scientist, other than Mendeleev, who contributed to the development of the modern periodic table of elements. Write a brief biography of this person and detail his or her scientific accomplishments.

78. The results and interpretations of chemistry experiments and studies are recorded and published in literally hundreds of scientific journals around the world. Visit the local library and look at several of the articles in a chemistry journal such as *The Journal of the American Chemical Society*. Write a brief summary of your observations regarding the format and style of writing in chemistry.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

79. What is chemistry? (Chapter 1)

80. What is mass? Weight? (Chapter 1)

81. Express the following in scientific notation. (Chapter 2)

a. 34 500
b. 2665
c. 0.9640
d. 789
e. 75 600
f. 0.002 189

82. Perform the following operations. (Chapter 2)

a. $10^7 \times 10^3$
b. $(1.4 \times 10^{-3}) \times (5.1 \times 10^{-5})$
c. $(2 \times 10^{-3}) \times (4 \times 10^5)$

83. Convert 65°C to Kelvins. (Chapter 2)

84. Graph the following data. What is the slope of the line? (Chapter 2)

Energy Released by Carbon

Mass (g)	Energy released (kJ)
1.00	33
2.00	66
3.00	99
4.00	132

STANDARDIZED TEST PRACTICE
CHAPTER 3

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Tables Use the table to answer questions 1 and 2.

Mass Analysis of Two Chlorine–Fluorine Compound Samples

Sample	Mass chlorine (g)	Mass fluorine (g)	%Cl	%F
I	13.022	6.978	65.11	34.89
II	5.753	9.248	?	?

1. What are the values for %Cl and %F, respectively, for Sample II?

- **a.** 0.622 and 61.65
- **b.** 61.65 and 38.35
- **c.** 38.35 and 0.622
- **d.** 38.35 and 61.65

2. Which of the following statements best describes the relationship between the two samples?

- **a.** The compound in Sample I is the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of definite proportions.
- **b.** The compound in Sample I is the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of multiple proportions.
- **c.** The compound in Sample I is not the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of proportions.
- **d.** The compound in Sample I is not the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of multiple proportions.

3. After elements A and B react to completion in a closed container, the ratio of masses of A and B in the container will be the same as before the reaction. This is true because of the law of _____ .

- **a.** definite proportions
- **b.** multiple proportions
- **c.** conservation of mass
- **d.** conservation of energy

4. All of the following are physical properties of table sugar (sucrose) EXCEPT _____ .

- **a.** forms solid crystals at room temperature
- **b.** appears as crystals white in color
- **c.** breaks down into carbon and water vapor when heated
- **d.** tastes sweet

5. A substance is said to be in the solid state if _____ .

- **a.** it is hard and rigid
- **b.** it can be compressed into a smaller volume
- **c.** it takes the shape of its container
- **d.** its matter particles are close together

6. Na, K, Li, and Cs all share very similar chemical properties. In the periodic table of elements, they most likely belong to the same _____ .

- **a.** row
- **b.** period
- **c.** group
- **d.** element

7. A heterogeneous mixture _______

- **a.** cannot be separated by physical means.
- **b.** is composed of distinct areas of composition.
- **c.** is also called a solution.
- **d.** has the same composition throughout.

8. The percent by mass of sulfur in sulfuric acid, H_2SO_4, is _____ .

- **a.** 32.69%
- **b.** 64.13%
- **c.** 16.31%
- **d.** 48.57%

9. Magnesium reacts explosively with oxygen to form magnesium oxide. All of the following are true of this reaction EXCEPT _______ .

- **a.** The mass of magnesium oxide produced equals the mass of magnesium consumed plus the mass of oxygen consumed.
- **b.** The reaction describes the formation of a new substance.
- **c.** The product of the reaction, magnesium oxide, is a chemical compound.
- **d.** Magnesium oxide has physical and chemical properties similar to both oxygen and magnesium.

10. Which of the following is NOT a chemical reaction?

- **a.** dissolution of sodium chloride in water
- **b.** combustion of gasoline
- **c.** fading of wallpaper by sunlight
- **d.** curdling of milk

THE PRINCETON REVIEW **TEST-TAKING TIP**

When Eliminating, Cross It Out Consider each answer choice individually and cross out choices you've eliminated. If you can't write in the test booklet, use the scratch paper. List the answer choice letters on the scratch paper and cross them out there. You'll save time and stop yourself from choosing an answer you've mentally eliminated.